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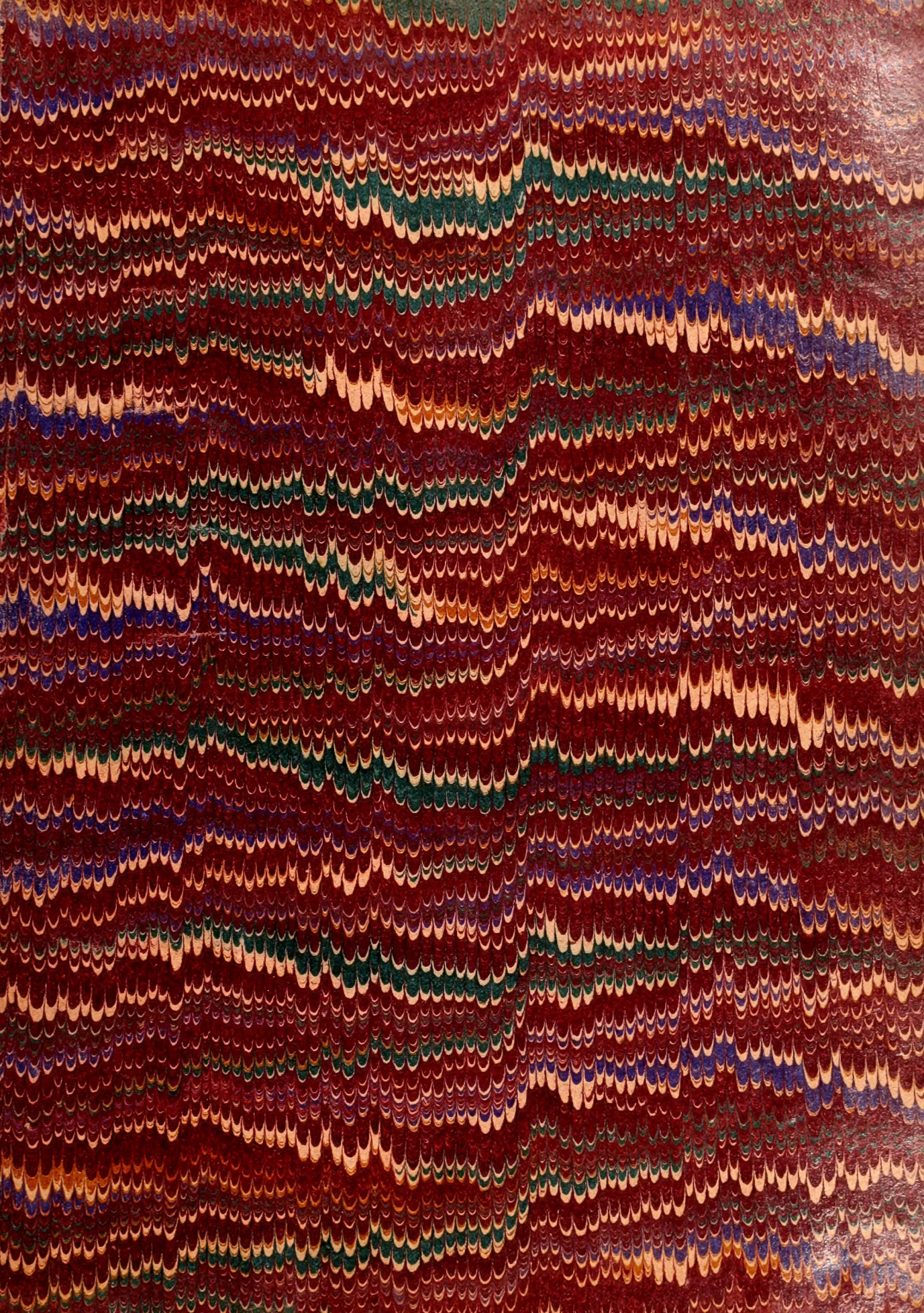

The Reduction of Permanganic Acid.

by Hydrogen and Ethylene

and

A Study of Some of Its Salts.

DISSERTATION

SUBMITTED TO THE BOARD OF UNIVERSITY STUDIES

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Austin M. Patterson.

1897.

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Acknowledgment

Some investigations in this paper were conducted at the computation laboratory under the auspices of Professor H. H. Aiken. To him I owe a great debt of gratitude. For the valuable advice of Professor Aiken, for the [illegible] the gift devised from his estate, and for instruction received from Doctor Matthews and Dr. Shattuck I am also grateful.

Prefatory Note.

Some explanations of the composition and character of this dissertation seem due. I had hoped to make a more complete study of the phenomena treated of in the First Part, but was unfortunately prevented from doing so by circumstances [illegible] which entitled the latter. [illegible] (Proc. Am. Philosophical Soc. 1891) was produced the [illegible] was [illegible] in which it stood that putting it in order for publication was out of the question for the time being. On Dr. Morse's advice the work described in the Second Part was then undertaken.

Part I.

The Reduction of Permanganic Acid by Hydrogen and Ethylene.

Introduction.

In 1896 Victor Meyer and Max von Recklinghausen published an article entitled, "The Gradual Oxidation of Hydrogen and Carbon Monoxide."[1] It had long been known that hydrogen is slowly attacked with oxidation by acidified solutions of potassium permanganate, and it was with the object of studying the course of this reaction that their experiments were made. But an unexpected phenomenon, the spontaneous liberation of oxygen, presented itself. Various theories as to its cause were considered, tested, and discarded, and the experimenters found themselves at a loss for an explanation.

Their attention was soon called, however, to some phenomena noted by Mond,

1 Ber. d. chem. Ges. 29, 2549

Hofman and Trachsler,[1] who had succeeded
in getting mercuric bromide and iodide to
evolve oxygen in connection with evolution of oxy-
gen. The manganese dioxide in the
two products formed when hydrogen
or sodium monoxide reacts with hydro-
gen peroxide, the suggestion
was made that in the presence the
oxidation of organic matter may be due.[2]

This question was considered in an
article by Wirth and Thomas,[3] who dis-
cussed the theories of direct
bearing and the problem. They assumed
that the American investigators had
experimented with neutral solutions
only, and that therefore the conditions
of experiment are not comparable.
They also regarded as an unwarrantable
objection the fact, that the oxida-

[1] Am. Chem. Jour. 18, 401 [2] Ber. d. chem. Ges. 30, 48
[3] " " " " 29, 2828

turn of oxygen which reacts during the absorptions, is much more rapid than that which takes place after the absorptions, of the gas. While this is the case with hydrogen and carbon monoxide it is not true, as will be seen from the present investigation, for the oxidizable gases.

The first difficulty was to get an apparatus free of the [illegible] recorded by Morse, Hopkins, and Walker[1] 35 mm with acid solutions for the second experiment. Morse and Hopkins[1], through carelessness, at first, overlooked it in their elaborate experiments. Failure was finally traced to rubber tubing which had come into use in the laboratory during the time that the gases had been in contact with it. Rubber tubing, even that of the highest grade, is a great mistake

[1] Am. Chem. Jour. 20, 721

activity. They advanced the hypothesis
hypothesis that the molecules of oxide
which are first formed have a strong
tendency to polymerize. That they
are able to incorporate the ferricyanic
acid, taking up more or less of it
and liberating the corresponding
It would then not be surprising, if,
as the concentration of the molecules in-
creased, their activity should vary.
Evidence in favor of this hypothesis
is the fact that large fluctuations must
be assigned to the ordinary constitut-
ive properties of precipitates by reason
of the minute proportions of water or
ions which they contain.

Contrary to the hypothesis of Haber
and Van Oordt-Vogelsang, the oxidations
were obtained by stirring with glass rods

More recently, Morse and Burer[1] have shown, that when carbon monoxide is absorbed by neutral solutions of ammonium chloride, there is no evolution of oxygen; but when acid is added, there is an evolution of oxygen which is due to that which is dissolved with the carbon monoxide. When the gas is absorbed by an acidified solution of ammonium chloride the same reaction takes place.

[illegible] the kind of product evolution is not connected with. The disappearance of the gas. Though the nature [illegible] to be, and that the reduction of oxygen takes place with-

[1] Am. Chem. Jour. 23. 313

a given time increases with increasing concentration of the hormone used.

These results lead to the conclusion that there is an interaction between the rate of absorption and the duration of exposure. For the longer of the times used, however, there appears to be a threshold of probability. In order to obtain additional evidence, it was thought advisable to study the phenomenon [illegible] under various conditions; the results of that study are now presented.

Description of the Apparatus.

The conditions of experiment were as
[illegible] possible. The [illegible] 12 cm. be-
[illegible] by [illegible]. Upon
these [illegible], [illegible] [illegible]
tubes were [illegible]. [illegible] [illegible] of
one hundred cubic centimetres, and pro-
vided with two stop-cocks — for the
introduction of liquids, the other for the
gas. The gas-burette was [illegible]
temperature was [illegible] by [illegible]
the [illegible] [illegible] which
[illegible] through [illegible]. The
temperature of the bath remained very
constant, at 25°, with a maximum
variation of about one degree.

The measure of the [illegible] attended
to the [illegible], [illegible], careful record
was kept of the number of [illegible]

performed by each Tube Varying each experiment. In an examination of the results of the experiments it may be noted that in any series where the only variant is the rate of revolutions, the maximum evolution of oxygen is shown by those tubes which revolve about thirty-six times per minute, a rate which allows the liquid to travel regularly from end to end. However, for slight variations in speed, these differences are so small that they may be neglected.

Preparation of Materials.

Hydrogen. - The hydrogen was generated electrolytically, and the ferricyanide by oxidation in alkaline solution. A resistance equivalent to twelve cubic centimetres at 0° and 760 mm was used

in each experiment.

Ethylene. - The ethylene formed was taken for the reducing gases of hydrogen, carbon monoxide, volume for volume, it was necessary, in order to have similar conditions, to dilute the pure hydrocarbon with four times its volume of air. Twelve cubic centimeters of this mixture were used in each experiment.

For preparing pure ethylene, advantage was taken of a reaction described by Wurtz[1], which takes place when ethylene bromide is brought in contact with the zinc-copper couple, water also being present:

$$C_2H_4Br_2 + Zn = ZnBr_2 + C_2H_4.$$

Ethylene bromide alone, even when heated in boiling water, reacts very slowly with zinc, owing to the formation of a film of zinc bromide on the surface of the metal.

[1] Jour. Chem. Soc. 27, 406

preventing further action. Cut off some wire, and coil one end, as shown, which will hold the bromide in solution. The reaction is not hindered in this way.

The generating apparatus is very simple and is recommended for the convenient preparation of the gas. In this a pure agate is chipped into bits of about the size of a kernel of corn, and the tube is broken out to a long capillary at both ends. Two pieces of stiff wire are bent and pushed about half way into the tube till the agate is covered and only a small air-space remains; the lower capillary is then closed by inserting it into a cork. Soon the reaction begins, and when the air in the upper capillary has been displaced, connection is made with the reducing

apparatus. Heat of reduction can be regulated by varying the proportion of alcohol in the mixture. One volume of alcohol and two of thallous bromide have been found to give [illegible] conduction, while if the proportions are reversed, the reaction is violent. The active electrode. It may be constructed by heating one end of an asbestos wire which is wound around the terminals. A small piece of platinum gauze has been soldered before the liquid substance itself, the wire removed and the ends [illegible] closed. The presence of the gauze keeps the liquid out of the hollow cylinder. The thallous hydroxide in this manner is entirely absorbed by hydrochloric acid.

In all the experiments the oxide resulting from the absorption of the hydrogen

[illegible] ethylene from the excess of ferrocyanide with the [illegible] value of one-tenth normal. Now twelve molecules of hydrogen are equivalent to one molecule of ethylene and to twelve molecules of potassium permanganate in the [illegible] solution. One cubic inch, the known ratio of twenty, to one, twenty, forms the [illegible] with the above [illegible] by the gram, and we [illegible]. [illegible] of the [illegible] required. [illegible] the [illegible] [illegible] with [illegible] experiment.

The [illegible] of potassium permanganate [illegible] by titrating the [illegible] solution [illegible] [illegible] [illegible]

of the tube, after evaporation, with the addition of sulfuric acid; the amount of oxygen liberated could then be calculated. In working with acid solutions the amount of acid added was denoted "3 K"; i.e., the equivalent of the hydrogen in the permanganate added by the pipet, plus three times the titration on the excess of the acid. In the dilute acid, the acid was not multiplied.

Experiments with Hydrogen.

Since [illegible] to the [illegible] solutions for the [illegible] with [illegible] solutions (Table I) and acid solutions (Table II). [illegible] remained in the tube at the end of each experiment and which might [illegible] in addition to the small amount of oxygen, was absorbed by [illegible]. was passed over spongy palladium heated to 100°. under these circumstances only hydrogen which has not been absorbed by the solution [illegible] the [illegible] of which [illegible] a [illegible] to [illegible] with [illegible] the [illegible] to [illegible] [illegible] to [illegible] [illegible] of the [illegible]

TABLE I. (NEUTRAL).

Time of Agitation in Hours.	Hydrogen unabsorbed C.c.	Oxygen liberated C.c.
24	[illegible]	0.00
24	[illegible]	[illegible]
45	0.7	[illegible]
45	0.6	[illegible]
47	1.0	[illegible]
47	0.5	[illegible]
48	0.6	[illegible]
53	1.4	0.00
72	0.5	0.00
72	[illegible]	[illegible]
72	[illegible]	0.36
72	[illegible]	[illegible]
74	[illegible]	[illegible]
90	[illegible]	[illegible]

At the end of twenty-four hours some hydrogen still remains, but in forty-eight hours most of it has been oxidized. The absorption of the remainder takes place with difficulty. These results and also others not here included, extending to periods of 216 hours, show that the last traces of the gas are not attacked readily, even by dilute alkali. However, for all practical purposes, we may assume that the absorption is complete by the end of twenty-four hours.

TABLE II (Acid)

Time of Agitation in Hours	Hydrogen unabsorbed C.c	Oxygen evolved C.c.	Percentage Reduction
4	0.7	5.78	4.82
4	0.7	—	—
4	0.6	3.88	3.23
4	0.2	6.21	5.17
4	0.0	4.66	3.88
4¼	0.0	6.35	5.29
4½	—	5.27	4.39
5	0.0	6.21	5.17
5	0.0	5.99	4.99
5	—	5.99	4.99
5	—	6.21	5.17

We see from the table that in acid solution the absorption is entirely complete before the end of the agitation. The evolution of the residual gas for hydrogen was more than

continued, and the tubes were centrated for
[illegible] periods with the object of determining
the amount of oxygen evolved. Table III shows
the results, duplicates being indicated by braces.

TABLE III. (ACID).

Hours of Agitation	Oxygen evolved.	Percent. Reductio
0	6.85	5.71
6	6.92	5.77
6	6.99	5.82
24	6.01	5.01
24	6.29	5.24
24	7.85	6.54
24	8.63	7.19
24	8.61	7.17
24	9.25	7.71
48	9.90	8.25
48	10.19	8.49
48	10.78	8.98
144	16.75	13.96
144	18.20	15.17

As has been stated, Morse and [illegible] found that only minute amounts of oxygen are evolved from neutral solutions in which [illegible] were observed; Table I shows the same to be true in the case of hydrogen. Also, Table II confirms the previous observations of acid evolution from acid solutions. It was important to know whether, when hydrogen is absorbed in neutral solutions and the solutions afterward acidified, oxygen would be liberated. Experiments were commenced with this end in view, and the results are recorded in Table III

TABLE IV. (NEUTRAL, THEN ACID).

Hours of Agitation		Oxygen evolved.	Percent. Reductic
Neutral	Acid	c.c.	
72	6	1.46	1.22
72	6	2.26	1.88
72	24	3.13	2.61
72	24	3.13	2.61
72	24	4.22	3.52
72	48	4.81	4.01
72	48	5.32	4.43

These series, made to run parallel, is incomplete. Experiments for still longer periods were in progress, but were abandoned, terminated by the fire [illegible].

These which were completed are sufficient, however, to show that liberation of oxygen does take place, though not to the same extent as in solutions acid from the start.

Experiments with Ethylene.

The action of ethylene on mercuric sulphate solutions of strengths, the action of ethylene on mercuric acetate, especially hydrocarbons, and hydrocarbons generally into aqueous mercuric sulphate solutions, has been studied by Hofmann and Wolf[1] in the following manner.

These authors, being interested mainly in the rate of absorption, contented themselves in the cases of ethylene and acetylene with the observation that they are oxidized almost instantaneously by permanganate solutions, and the attendant phenomena were left unstudied by them. Table I records the results of the action of ethylene upon the neutral solution. In the first few experiments the residual gas was tested for its hydrocarbons by fuming it over pyrosulphuric acid, but no unabsorbed ethylene could be detected at the end of each

[1] Ber. d. chem. Ges. 30, 1935

the shortest period of agitation.

TABLE V. (NEUTRAL)

Time of Agitation Hours.	Apparent Reduction C.c. of Oxygen.
½	−1.44
1	−0.87
2	0.00
3	0.22
5	0.22
5	0.22
5	0.29
16⅔	0.57
70	0.36
70	0.36

Although the ethylene is so quickly absorbed, the titrations indicate that its oxidation to carbon dioxide and water is not completed until after one agitations of [illegible]

[illegible] to be subtracted. [illegible] titrated [illegible] acid. [illegible] acid, and the [illegible] oxalic[1] [illegible] solution [illegible] permanganate.

In the first experiments with acid solutions (recorded in Table II) the residual [illegible] Total [illegible], but [illegible] could be titrated at the end of fifteen minutes, the constant [illegible] Tables II and III. The subtractions are connected by braces; the letters [illegible] the order of the [illegible].

[1] Ber. d. chem. Ges. 21. 1230.

TABLE VI. (ACID).

	Hours of Agitation	Oxygen evolved	Per cent. Reduction
	¼	−0.07	
	¼	0.00	
	½	0.29	
	½	0.57	
a	24	0.79	0.66
a	24	1.01	0.84
b	24	1.16	0.97
b	24	1.81	1.50
c	48	2.02	1.68
d	72	4.55	3.79
d	72	5.85	4.87
e	72	5.13	4.27
e	72	5.64	4.70

[illegible] Table I that with respect to the [illegible] of oxygen from neutral solutions by [illegible] of the [illegible] Table VII, according to

Table IV gives hydrogen peroxide solutions which contained sodium hydroxide and were neutralized with sulphuric acid before agitation, while Table V solutions were then acidified:

TABLE VII. (NEUTRAL, THEN ACID).

	Hours of Agitation		Oxygen evolved	Percentage
	Neutral	Acid	C.c.	Reduction.
A	3	24	1.64	1.32
	3	24	1.62	1.27
B	3	24	1.84	1.44
	3	24	2.46	2.05
C	3	48	3.41	2.65
	3	48	3.18	2.46
D	3	72	7.79	6.11
	3	72	10.26	7.81
E	3	72	6.39	5.61
	3	72	7.44	6.40
F	3	72	3.97	3.11
	3	72	11.48	9.13

* Acid used, dilute. Time for acid not included in the above.

Here again it is to be regretted that the experiments could not be extended to longer periods of time.

Tables of Averages.

In order that the variations of the foregoing results may be more easily seen, the averages of all the experiments in the same unit are given as follows:

A. Total Oxygen Evolved.

Hydrogen.			Ethylene.		
Time in Hours	Acid (Table III)	Neutral, then Acid (Table IV)	Time in Hours	Acid (Table VII)	Neutral, then Acid (Table VIII)
6	6.92 cc.	1.86 cc.	24	1.19 cc	1.82
24	7.78	3.41	48	2.02	3.18
48	10.29	5.06	72	5.29	8.04
144	17.47	—			

B. Oxygen Evolved in Successive Periods.

Hours	(Table III)	(Table IV)	Hours	(Table VI)	(Table VII)
1–6	6.92 cc.	1.86 cc.	1–24	1.19 cc.	1.82 cc.
7–24	0.86	1.63	25–48	0.83	1.36
25–48	2.51	1.57	49–72	3.27	4.86
49–144	6.18	—			

C. Rates of Evolution (c.c. per hour).

Hours	(Table III)	(Table IV)	Hours	(Table VI)	(Table VII)
1–6	1.15	0.31	1–24	0.05	0.08
7–24	0.05	0.09	25–48	0.035	0.06
25–48	0.10	0.065	49–72	0.14	0.20
49–144	0.07	—			

Conclusions.

Palladium precipitates ... different ... have thus been studied:

(1) By hydrogen in acid solution

(2) By hydrogen in neutral solution

(3) By ethylene in acid solution.

(4) By ethylene in neutral solution.

... are comparable to those prepared by Morse and Byers:

(5) By carbon monoxide in acid solution.

(6) By carbon monoxide in neutral solution.

... these ... exhibit distinct characteristics ... the reduction of hydrogen from acidified permanganate solutions in which they are suspended.

1. ... previously noticed in the case of carbon monoxide, the oxides formed by ethylene and by gases in neutral solutions ... and ... reduction ... that

conductivity. They were not affected by the addition of acid.

2. The longest [illegible] of [illegible] was that of [illegible] in which the resistance [illegible] in acid solutions. The [illegible] is a rapid [illegible] rate of attack by these acids [illegible] [illegible] and this is [illegible] [illegible] by a rapid [illegible] [illegible]; the [illegible] conductivity, [illegible] [illegible]. The [illegible] [illegible] [illegible] definite limit to the [illegible]. At the end of the longest period studied, which is six days.

3. [illegible] [illegible] [illegible] [illegible] additional work [illegible] [illegible], which, [illegible] [illegible] [illegible] [illegible] [illegible]. [illegible] [illegible] [illegible] [illegible] [illegible] [illegible] [illegible] [illegible] [illegible] [illegible] [illegible].

4. [illegible] [illegible] [illegible] [illegible] [illegible] [illegible] [illegible] [illegible] [illegible] [illegible]

[illegible] difference between the oscillation of the spring and the point at which the maximum rate of vibration of oxygen is attained.

The results summarized in the first and second conclusions are valuable as confirmatory evidence, but possess no features that are new. The facts upon which the third is based admit of [illegible] the views of Morse and Reese to be true. The [illegible] of the [illegible] [illegible] [illegible]

of the hydrogen is absorbed during the first twenty-four hours, but not seventy-two hours are required to complete the oxidation. Consequently, during the last forty-eight hours the oxidizing part of the cycle is carried on by the oxygen-carriers and by the air. The amount of oxygen-carrier used is [illegible] the [illegible] to [illegible] the [illegible] [illegible] [illegible]

monoxide is allowed to stand quietly over
an acidified solution of ferrous sulphate,
it [illegible] with the oxidation of the iron.

The oxide precipitated by the latter in-
variably contains cuprous oxide, as can be
seen from the [illegible] with the
amount left in solution. This, and
many other facts that could be given,
show conclusively that cuprous oxide
[illegible] by the hydrolytic [illegible]
[illegible] with the salts of copper,
and not with the oxide precipitated.
The exact [illegible] the concentration of
the [illegible] at which it is possible
to [illegible] at which the rate of
[illegible] is reduced, [illegible]
much is sure, that it varies many hundred
[illegible] the [illegible] concentration
[illegible] of the [illegible]

It appears possible to eliminate at least
a few basic complications, i.e., the presence
of known, and number of unfortunate events
presumably be subtracted from unfortunate
incidents within [illegible].

Part II.

Some Salts of Permanganic Acid

Solubilities of Permanganates of the Alkali Metals.

Permanganic acid and some permanganates were first prepared in 1859 by Mitscherlich[1] whose method consists in decomposing potassium permanganate with sulphuric acid and in crystallizing the permanganic acid in dark red crystals which decompose at 100°, "diffusing violet vapors and oxygen gas, and leaving behind manganese dioxide". Permanganic acid is also formed by the action of concentrated sulphuric acid on potassium chlorate, perchloric acid being formed at the same time. Permanganic acid is also formed by the action of concentrated sulphuric acid on potassium chlorate, perchloric acid being formed at the same time.

Permanganic acid and its salts, especially the latter, have been investigated in this laboratory, have been found to be of extraordinary stability, the slight solubility of the corresponding salts of the alkali metals becomes interesting by way

[1] Ber. d. chem. Ges. 26. 1016

of contact. The remaining [illegible] solutions
the following method was employed. The
[illegible] bottle, containing [illegible]
and an excess of the recrystallized
salt was placed in water-bath at a constant
temperature. The bottle was allowed to re-
main in the bath for several hours, and
was agitated thoroughly at short intervals.
When saturated the concentration of
the saturated solutions was
determined, and the latter were then drawn
by suction through a glass tube containing
a plug of cotton. The first portions
of the filtrate were discarded. The strength
of the filtered solutions were determined
by titration against the accurately standard-
ized solutions of oxalic acid, a method
whose error is smaller than that encount-
ered in reading the thermometer.

Care was taken to keep the temperature of the following operations the same, or as little unlike as that of the mixture in order to prevent any of the salt from crystallizing out.

Volumetric Measurements. Twenty grams of molybdenum concentrate, chosen by specimen analysis to be of great purity, was added to the equivalent amount of a mixture of ferric oxide and ferrous oxide in the ratio of Fe3O4. Oxide was then dissolved in a minimum of acid to the sulfate condition. The ferrous to ferric ratio of the mixture was then determined by titration with potassium permanganate, although as much as 10 per cent in excess could be tolerated by the ferric salt. Of the salt thus formed, the ferrous part was then titrated with

Calculated for $RbMnO_4$	58.21
Found	58.15

At 2°, 100 cubic centimeters of solution contain 0.46 grams of $RbMnO_4$.

At 60° 100 cubic centimeters contain 4.68 grams.

At 59° 100 cubic centimeters contain 1.25 grams.

Potassium Permanganate. - The [illegible] according to whom 100 parts of water at 15° dissolve 6.25 parts of potassium [illegible].

[illegible] to be able to construct curves for all these salts, we conducted the [illegible] at two other temperatures, [illegible] the results are

At 0° 100 cubic centimeters of solution contain 2.84 grams of $KMnO_4$.

At 30° 100 cubic centimeters contain 8.69 [illegible]

At 15° 100 cubic centimeters of solution contain 5.22 grams of $KMnO_4$.

On repetition it was found that:

At 15.3° 100 cubic centimeters contain 5.30 grams.

1.035. Translated into the terms used by us, the commonly recorded value reads:

[illegible]

So our best [illegible]

[illegible]

[illegible]

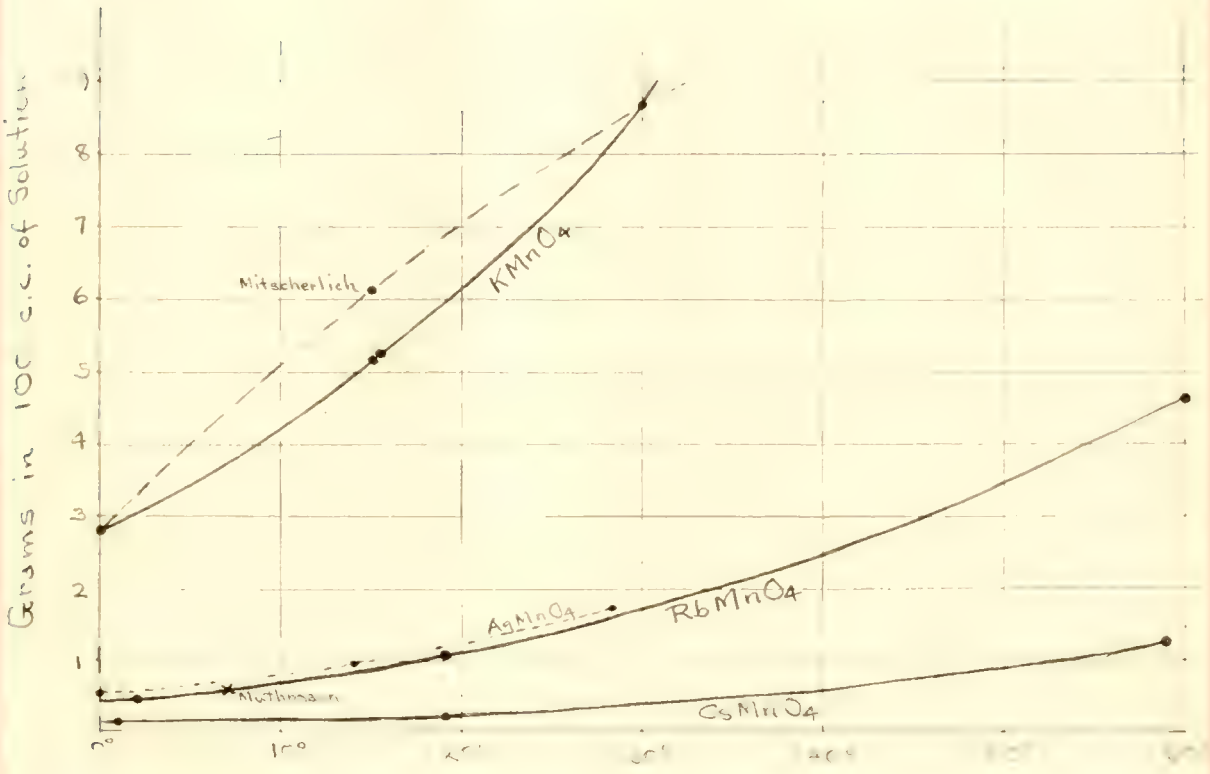
[

[illegible]

[illegible]

[illegible]

Solubility Curves.



Chromic Permanganate.

Some chromic oxide, prepared by [illegible] was dissolved in permanganic acid, and a portion of the liquid was treated with hydrogen peroxide and a little sulphuric acid. It was expected that in this way the permanganate would be reduced, while any chromic acid formed on oxidation would be further oxidized to the perchromic condition. On shaking with ether the ether was seen to be colored green while the aqueous solution remained colorless. The ether, [illegible] portion treated in the same way gave an intensely blue ethereal solution. After a few minutes standing the ethereal solution was poured into [illegible] potassium permanganate, dilute sulphuric acid, hydrogen peroxide and water. The ether solution [illegible] water [illegible]

Nitric acid is known by the name aqua fortis. Commercial nitric acid is usually a solution, but it is very unstable and easily decomposed by the action of light, evaporation, oxides, and organics.

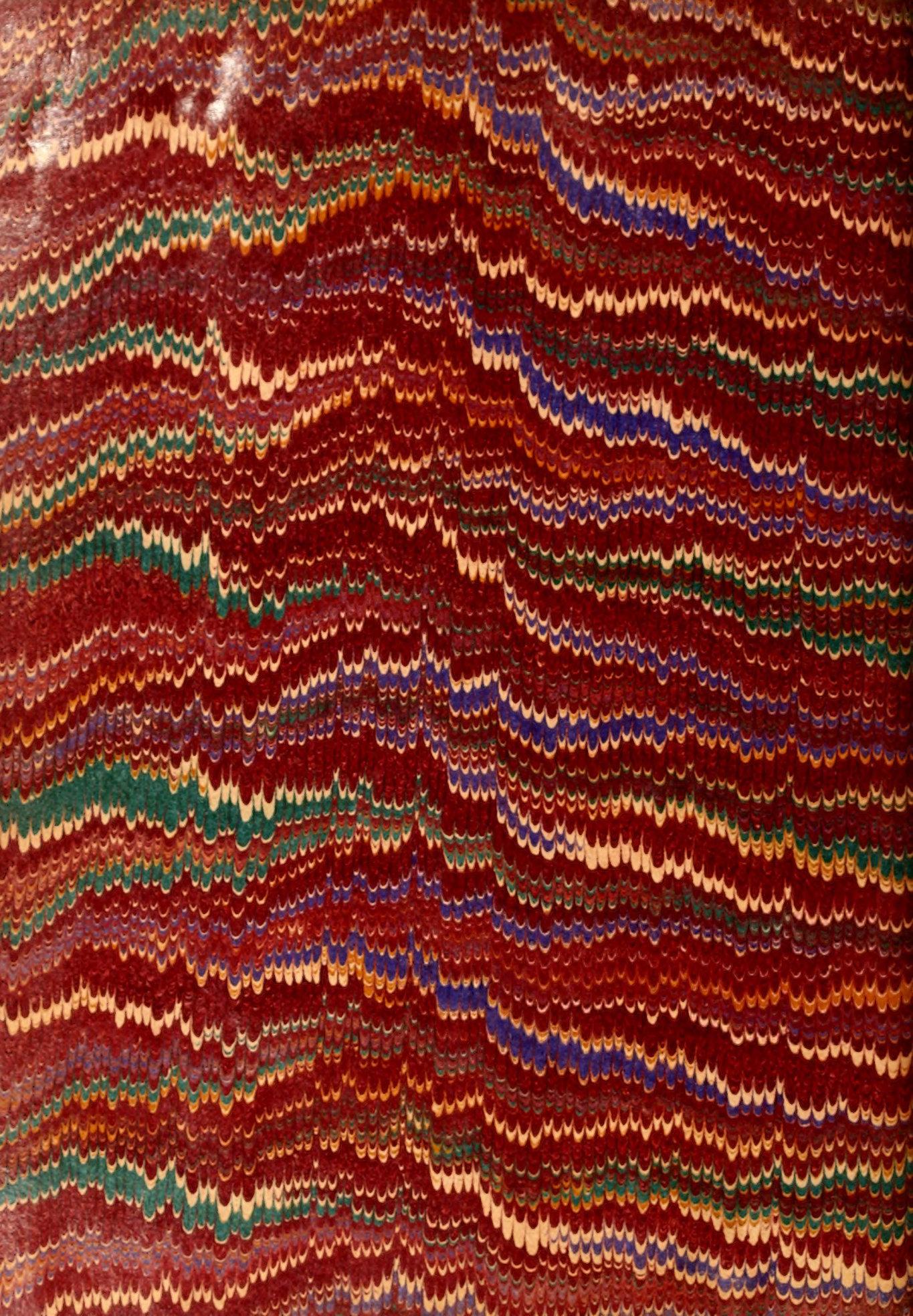
Lead Permanganate.

According to Fernekorn, lead permanganate is precipitated from a solution of lead nitrate by the addition of potassium permanganate, but is not known to exist, probably due to nitric acid. Nitrates are not in particular of interest, but many that would exist in sunlight by permanganate oxidation.

Experiment. - Obtain a liter of permanganate and solution, containing 7% excess of acid was neutralized by ammonia with an excess of lead oxide, and shaken with powder

[illegible] with [illegible] with the [illegible]
with [illegible] with [illegible]. It may
be said in passing that this method has
recommended itself as the most [illegible]
trouble for the detection of small amounts
of [illegible] in concentrated [illegible]-
ric solutions. The solutions used were
filtered, if necessary, to remove [illegible]
of precipitates and the reaction with [illegible]
carbonates with a stream of [illegible]
at about 40° was substituted. In this
way [illegible] could be [illegible],
without [illegible], but after cooling the
concentration the solution began to de-
posit large quantities of a brown substance
resembling [illegible], but con-
taining less [illegible]. It is
necessary, [illegible] of [illegible]
[illegible] of [illegible] of [illegible]

was neutralized with freshly precipitated copper oxide which on account of the reaction was observed with dry potassium carbonate. The solution was then filtered and by very slow heating on the water-bath was evaporated to about its one-eighth its volume without much decomposition. As soon as oxygen began to be given off, the solution was again filtered and transferred to a bulb-jar containing sulphuric acid. On further evaporation irregular crystals of a salt were deposited, which is probably soluble in a small amount of water, and on it to all the reactions of sulphuric ferrous sulphate. It belongs to the monoclinic system. Distinct evidences of crystal form have been observed, but no good crystals have as yet been obtained. Specific-gravity determinations which are necessary for

Biographical Sketch.

Lincoln M. Patterson was born in Damascus, Syria, May 31, 1876. He attended the public schools of Xenia, Ohio, and prepared for college at a private school in the same city. In 1893 he entered Princeton College, and in 1897 graduated with the degree of A.B. from Princeton University. In the fall of 1897 he entered the Ogden Graduate University, where he has since pursued the study of chemistry, mineralogy, and geology.